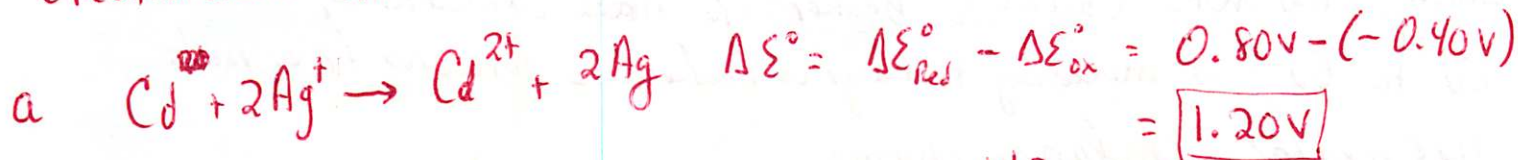


Electrochem HWwk

1



b $\Delta G^\circ = -nFE^\circ = -(2)(96,485\text{C/mol})(1.20\text{V}) \left(\frac{1\text{kJ}}{1000\text{J}}\right) = \boxed{-231\text{ kJ/mol}}$

b $\log K = \frac{nE^\circ}{0.0592} = \frac{(2)(1.20\text{V})}{0.0592}$ $\log K = 40.5$ $K = \boxed{3.47 \times 10^{40}}$

d $E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0592}{n} \log Q$

New $[\text{Cd}^{2+}] = 1\text{M} \times 100\text{mL} = 100\text{mmol}$
 $50.0\text{mL} \times 2 = 100\text{mmol}$
 $\frac{100\text{mmol}}{200\text{mmol}} = \frac{100\text{mmol}}{150\text{mL}} = \boxed{1.33\text{M Cd}^{2+}}$

$E_{\text{cell}} = 1.20\text{V} - \left(\frac{0.0592}{2}\right) \log \frac{1.33\text{M}}{1.0\text{M}} = \boxed{1.19\text{V}}$ or 1.20V

e If Ag^+ ion decreases $\text{Cd}^{2+} \uparrow$

$0.40\text{M} \times \frac{1\text{mol Cd}^{2+}}{2\text{mol Ag}^+} = 0.20\text{M}$

$E_{\text{cell}} = 1.20\text{V} - \frac{0.0592}{2} \log \frac{1.20\text{M}}{0.60\text{M}} = 1.19\text{V}$

f to keep the charge in each beaker neutral by allowing ions to move into solutions in each beaker to balance charge, it also completes the circuit.

g Ag metal \rightarrow cathode

Cd metal \rightarrow anode

h e^- flow toward Ag metal

i) anions flow into $\text{Cd}(\text{NO}_3)_2$ beaker b/c more oxidation of Cd to Cd^{2+} is increasing + charge in beaker so anions flow into this beaker to balance charge.

j) Cell voltage will increase

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0592}{n} \log \frac{[\text{Cd}^{2+}]}{[\text{Ag}^+]}$$

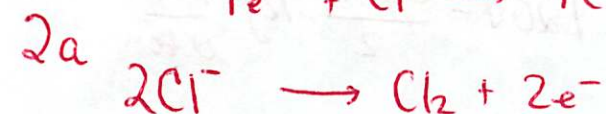
if $\text{Ag}^+ \uparrow$ the $\log Q$ will be (-) so voltage gets added to E_{cell}° and $E_{\text{cell}} \uparrow$

k) Cell voltage will decrease b/c $[\text{Ag}^+] \downarrow$ w/ ppt of AgCl

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0592}{n} \log \frac{[\text{Cd}^{2+}]}{[\text{Ag}^+]}$$

now $\log Q$ will be (+) so voltage gets subtracted from E_{cell}° and $E_{\text{cell}} \downarrow$

l) No change b/c $\log Q$ will stay the same and b/c $[]$ of neither ion will change and $\log Q$ will be 0 and $E_{\text{cell}} = E_{\text{cell}}^{\circ}$.



$$b \quad 2.00 \text{ hr} \times \frac{60 \text{ min}}{1 \text{ hr}} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{0.250 \text{ C}}{1 \text{ s}} \times \frac{1 \text{ mole}}{96500 \text{ C}} = 0.0186 \text{ mole } e^-$$

$$0.521 \text{ g Fe} \times \frac{1 \text{ mole Fe}}{55.85 \text{ g Fe}} = 0.00932 \text{ mol Fe}$$

$$\frac{0.0186 \text{ mole } e^-}{0.00932 \text{ mol Fe}} = \frac{2 \text{ mole } e^-}{1 \text{ mol Fe}} \rightarrow \text{Fe}^{2+}$$

so $\boxed{\text{FeCl}_2}$



d $0.00932 \text{ mol Fe} \times \frac{1 \text{ mol Cl}_2}{1 \text{ mol Fe}} = 0.00932 \text{ mol Cl}_2$

$$V = \frac{(0.00932 \text{ mol})(0.08206 \frac{\text{L atm}}{\text{mol K}})(25^\circ\text{C} + 273)}{\frac{750 \text{ mm Hg}}{760 \text{ mm Hg}}}$$

$$= 0.236 \text{ L Cl}_2$$

$$e \quad 3.00 \text{ g Cl}_2 \times \frac{1 \text{ mol Cl}_2}{71.0 \text{ g Cl}_2} \times \frac{2 \text{ mole}^-}{1 \text{ mol Cl}_2} \times \frac{96500 \text{ C}}{1 \text{ mole}^-} = \frac{8155 \text{ C}}{3600 \text{ s}} = 2.27 \frac{\text{C}}{\text{s}} = \boxed{2.27 \text{ A}}$$

$$1 \text{ hr} \times \frac{3600 \text{ s}}{1 \text{ hr}} = 3600 \text{ s}$$